

Chapter-8

Lesson-1

I – Chemical Bonding

A **chemical bond** is a strong attraction between two atoms and/or ions. There are three general types of chemical bonds: **ionic**, **covalent**, and **metallic**. All chemical bonding involves electrons.

A) The Octet Rule:

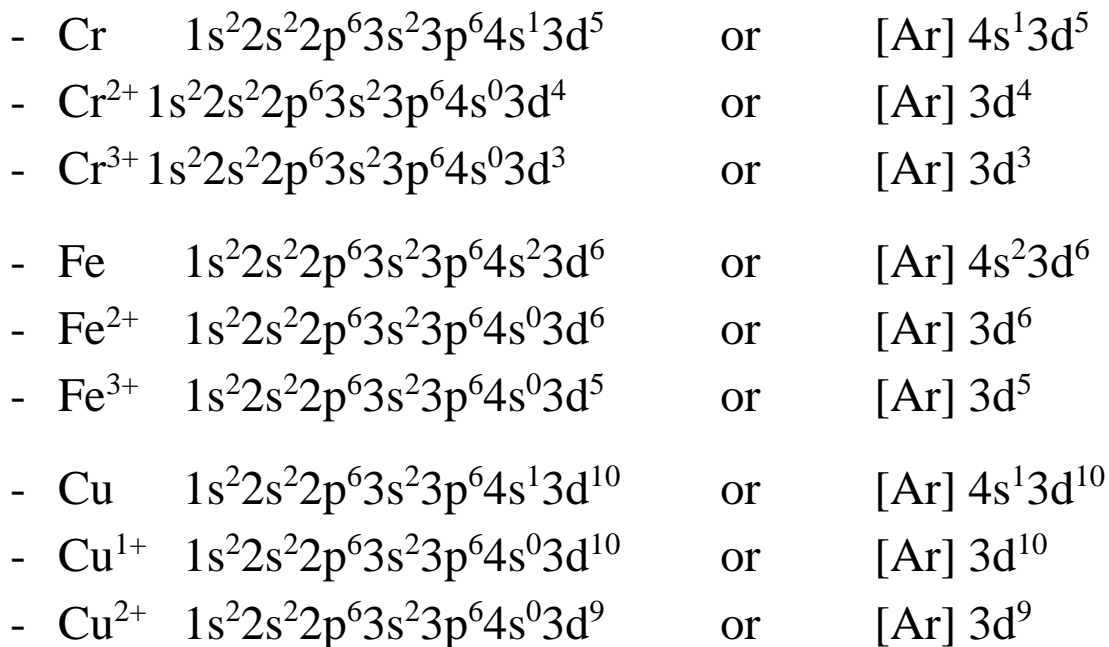
- Atoms of noble gases are extremely stable and all atoms of noble gases have a full valence which (except for helium) constitutes eight electrons. This stable eight valence electron arrangement is called a **stable octet**.
- Atoms will tend to gain, lose, or share electrons in order to achieve the stable octet electron configuration of the nearest noble gas. This is called the **octet rule**.

B) Electron Configurations of Monoatomic Ions:

- **Positive monoatomic ions** of the main group elements [*groups 1, 2, and 13 through 18*] will lose all their valence electrons in order to achieve a stable noble gas electron configuration.
 - K $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 \rightarrow K^{1+} 1s^2 2s^2 2p^6 3s^2 3p^6$ or [Ar]
 - Ca $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 \rightarrow Ca^{2+} 1s^2 2s^2 2p^6 3s^2 3p^6$ or [Ar]
 - Al $1s^2 2s^2 2p^6 3s^2 3p^1 \rightarrow Al^{3+} 1s^2 2s^2 2p^6$ or [Ne]
- **Negative monoatomic ions** will gain as many electrons as needed to fill the valence in order to achieve a stable noble gas electron configuration.
 - Cl $1s^2 2s^2 2p^6 3s^2 3p^5 \rightarrow Cl^{1-} 1s^2 2s^2 2p^6 3s^2 3p^6$ or [Ar]
 - S $1s^2 2s^2 2p^6 3s^2 3p^4 \rightarrow S^{2-} 1s^2 2s^2 2p^6 3s^2 3p^6$ or [Ar]
 - N $1s^2 2s^2 2p^3 \rightarrow N^{3-} 1s^2 2s^2 2p^6$ or [Ne]

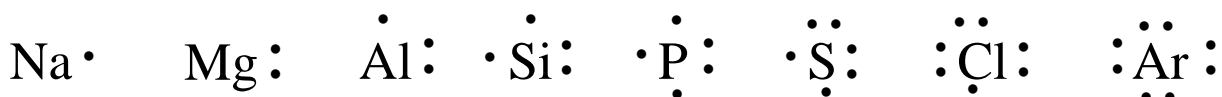
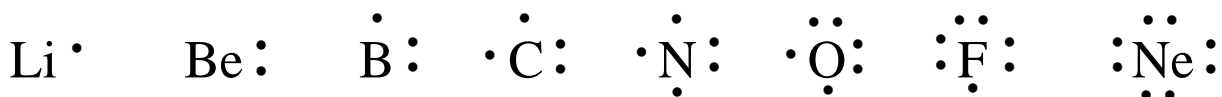
Monoatomic ions are isoelectronic with the nearest noble gas.

- **Positive monoatomic ions** of the transition elements will lose all their valence electrons and often high energy core electrons in order to achieve a stable noble gas electron configuration



C) Lewis Models for Atoms and Monoatomic Ions:

- Usually, only valence electrons are involved in bonding.
- In 1916, G.N. Lewis proposed a model to show the valence electrons in atoms and to follow their actions during bond formation.
- This model is called the **Lewis electron-dot model** or simply the **Lewis model**.



D) Sizes of Monoatomic Ions:

- Cations are smaller than their neutral atoms.

Na (0.186 nm atomic radius)

Na¹⁺ (0.095 nm ionic radius)

WHY?

Cations shed outer electrons and the remaining inner (core) electrons are attracted more by the protons in the nucleus.

- Anions are larger than their neutral atoms.

Cl (0.099 nm atomic radius)

Cl¹⁻ (0.181 nm ionic radius)

WHY?

Anions gain outer electrons and the additional outer electrons add to the repulsion between all outer electrons.

II – Ionic Bonding

- An **ionic bond** is an electrostatic attraction between ions of opposite charge. (BONDING BETWEEN IONS!)
- Simple ionic compounds generally result from the interaction of metals and nonmetals.
- The ions in an ionic bond may be formed from atoms ***by the transfer*** of one or more electrons from one atom to another.

A) Formulas of Ionic Compounds:

- binary ionic compounds are composed of positive metallic monoatomic ions and negative nonmetallic monoatomic ions.

NaCl

Cu₂S

Fe₂O₃

- Ternary and quaternary ionic compounds are composed of...
 - (i) positive metallic monoatomic ions and negative PAI's.
 - (ii) positive PAI's and negative nonmetallic monoatomic ions.
 - (iii) positive PAI's and negative PAI's.

CuSO₄

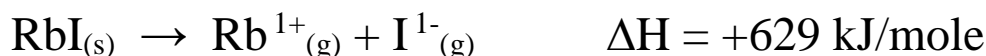
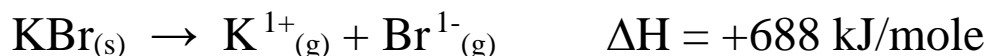
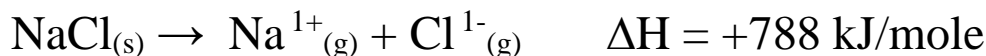
NH₄Cl

KH₂PO₄

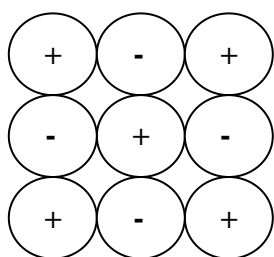
- *The bonding between the atoms within a PAI is always covalent*

B) Lattice Energies of Ionic Compounds:

- **Lattice energy** is the energy **ABSORBED** by an ionic compound in order to be decomposed into its gaseous cations and anions.
- The **lattice energy** for an ionic compound is **always** positive.
- The larger the lattice energy is the more stable and the stronger the ionic bond is.

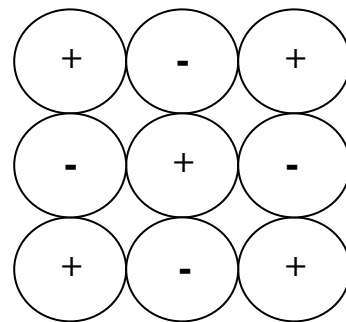


- A compound of a small cation and a small anion has a large lattice energy value whereas a compound of a large cation and a large anion has a lesser lattice energy value.

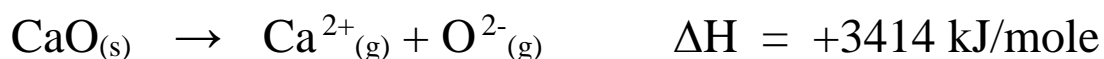
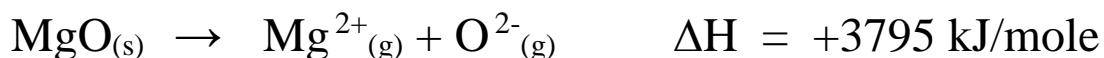
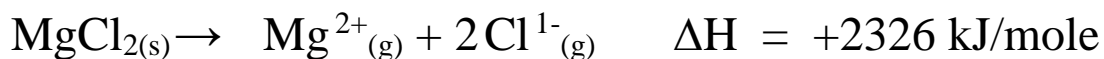
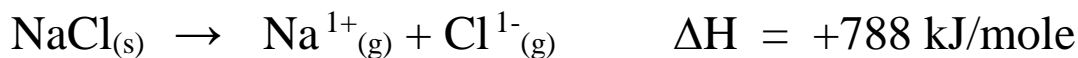


Small ions can get closer than larger ions and thus the ionic bond is stronger (larger lattice energy)!

$$E_{el} = \frac{kQ_1Q_2}{d}$$



- A compound's lattice energy is also dependent on the charges of the ions. The greater the ionic charge, the greater the lattice energy will be.



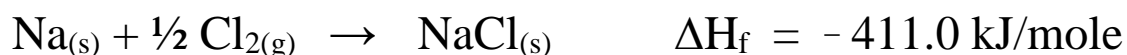
- For a given arrangement of ions, the lattice energy increases as (i) the charges on the ions increase and (ii) their radii decrease.

$$E_{el} = \frac{kQ_1Q_2}{d} \quad \text{where } d \approx r_{(+)\text{ion}} + r_{(-)\text{ion}}$$

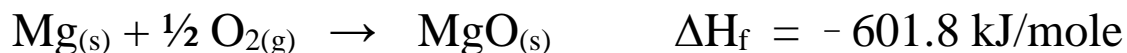
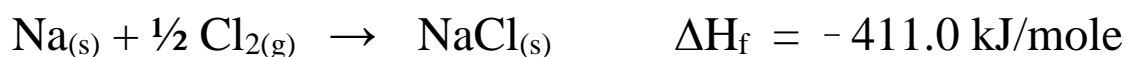
C) Heats of Formation of Ionic Compounds:

- **Heat of formation** is the energy absorbed or released when a compound is formed from its free elements.
- The heat of formation for an ionic compound is **always** negative.

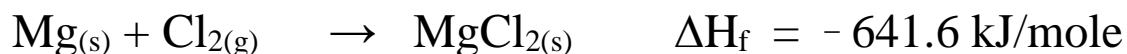
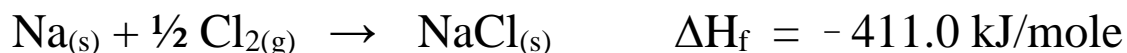
- **Compounds with the same ionic charges but different ionic radii**



- **Compounds with different ionic charges but similar ionic radii**



- **Compounds with different cations but the same anion**



Q1: Write the complete electron configurations for the following ions: (i) K^{1+} , (ii) Ca^{2+} , (iii) Sc^{3+} , (iv) F^{1-} , (v) O^{2-} , and (vi) N^{3-}

A1: (i) $1s^2 2s^2 2p^6 3s^2 3p^6$ (iv) $1s^2 2s^2 2p^6$

(ii) $1s^2 2s^2 2p^6 3s^2 3p^6$ (v) $1s^2 2s^2 2p^6$

(iii) $1s^2 2s^2 2p^6 3s^2 3p^6$ (vi) $1s^2 2s^2 2p^6$

Q2: Rank the following ions from smallest radius to largest radius: K^{1+} , Ca^{2+} , and Sc^{3+} ,

A2: $\text{Sc}^{3+} < \text{Ca}^{2+} < \text{K}^{1+}$

Q3: Rank the following ions from smallest radius to largest radius: F^{1-} , O^{2-} , and N^{3-}

A3: $\text{F}^{1-} < \text{O}^{2-} < \text{N}^{3-}$

Q4: Arrange the following ionic compounds in order of increasing melting point and explain why: CaO, KF, and ScN

A4: $\text{KF} < \text{CaO} < \text{ScN}$

All of the ions have the same electron configuration and thus the distance between the ions will be similar. However, the ionic charges are different: in ScN both ions have a 3+ charge, in CaO both ions have a 2+ charge, and in KF both ions have a 1+ charge. The larger the product of the charges, that greater the lattice energy, and thus the greater the melting point will be.

Q5: Write the complete electron configurations for the following ions: (i) Co^{3+} and (ii) Ti^{4+}

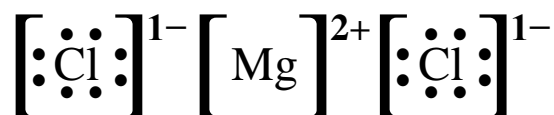
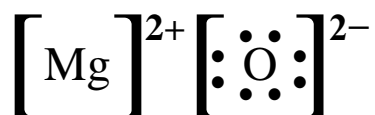
A5: (i) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$ (ii) $1s^2 2s^2 2p^6 3s^2 3p^6$

Q6: Write the Lewis models for the following atoms or ions: (i) In^{3+} , (ii) Kr, (iii) Te^{2-} , (iv) Co^{2+} , (v) As, and (vi) I^{1-}

A6: (i) $\left[\text{In} \right]^{3+}$ (iv) $\left[\text{Co} \right]^{2+}$
(ii) $:\ddot{\text{Kr}}:$ (v) $\cdot\ddot{\text{As}}\cdot$
(iii) $\left[:\ddot{\text{Te}}: \right]^{2-}$ (vi) $\left[:\ddot{\text{I}}: \right]^{1-}$

Q7: Write the Lewis models for the following ionic compounds: (i) magnesium oxide and (ii) magnesium chloride

A7: (i) magnesium oxide (ii) magnesium chloride



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Lesson-2

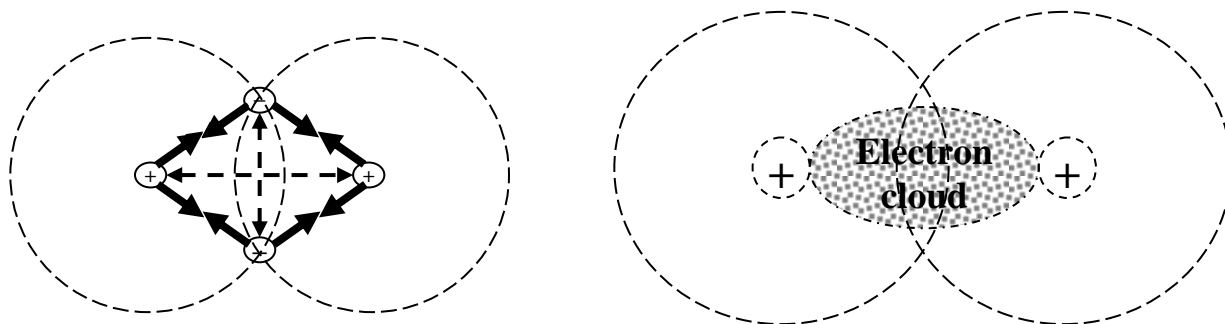
III – Covalent Bonding

- A covalent bond consists of a pair of electrons that are simultaneously attracted by two adjacent nuclei.
- A covalent bond consists of a pair of shared electrons between two nonmetal atoms in a binary compound.
- A covalent bond consists of a pair of shared electrons between any two atoms in a polyatomic ion.
- A covalent bond consists of a pair of shared electrons between a hydrogen atom and a (-) polyatomic ion.

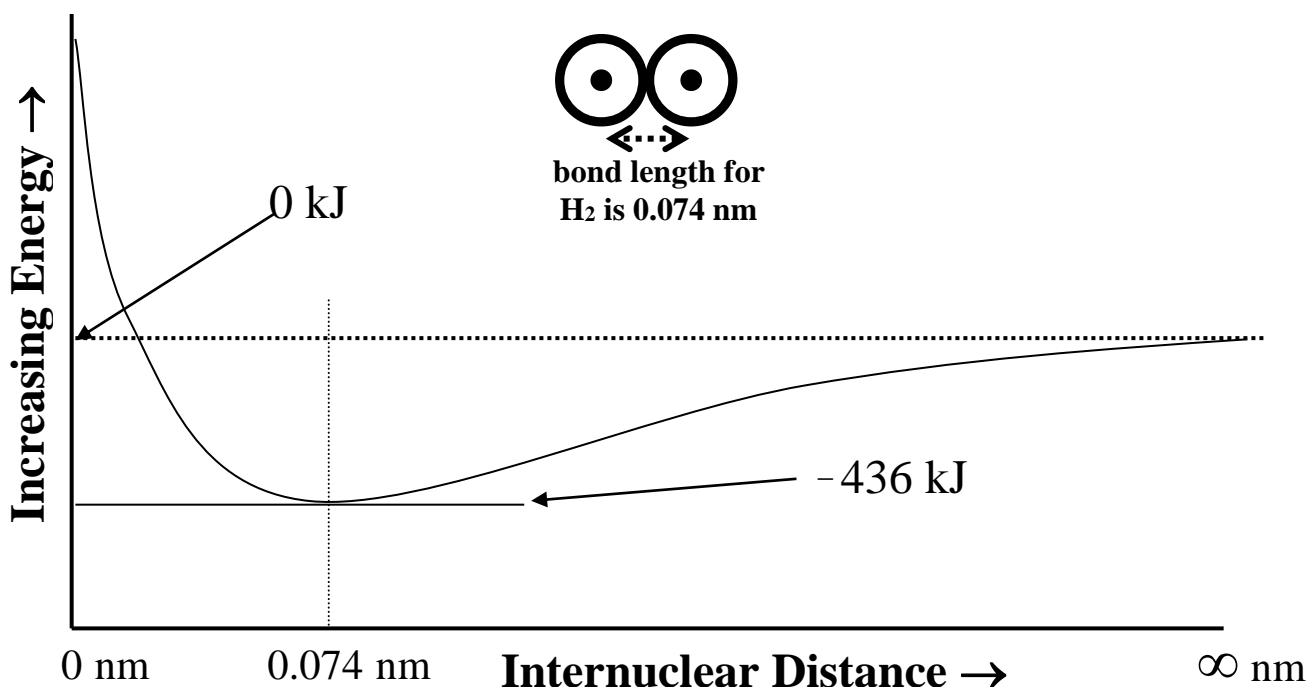
Q1: Why are two nuclei more stable when they share a pair of electrons?

A1: In 1927, W. A. Heitler & T. London used quantum mechanics to exactly solve this problem for a hydrogen molecule.

As two hydrogen atoms approach each other, an attractive force begins to grow between the atoms as a result of the simultaneous attraction the nuclei have for both electrons.



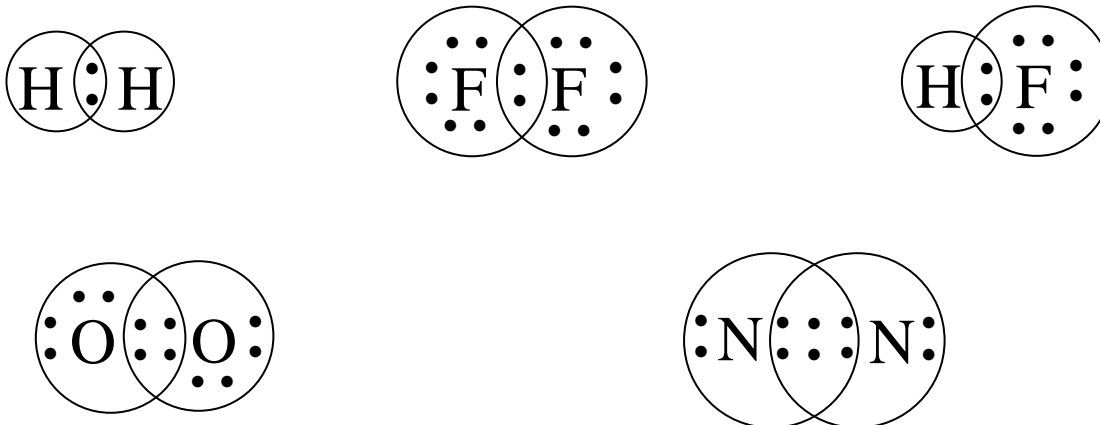
Internuclear and inter-electron forces of repulsion also begin to grow but the attractive forces outweigh the repulsive forces and a net attractive force results. Quantum mechanical calculations show that the electron density is heaviest between the two nuclei. Thus the two nuclei are electrostatically attracted to the heavy concentration of negative charge found between them.

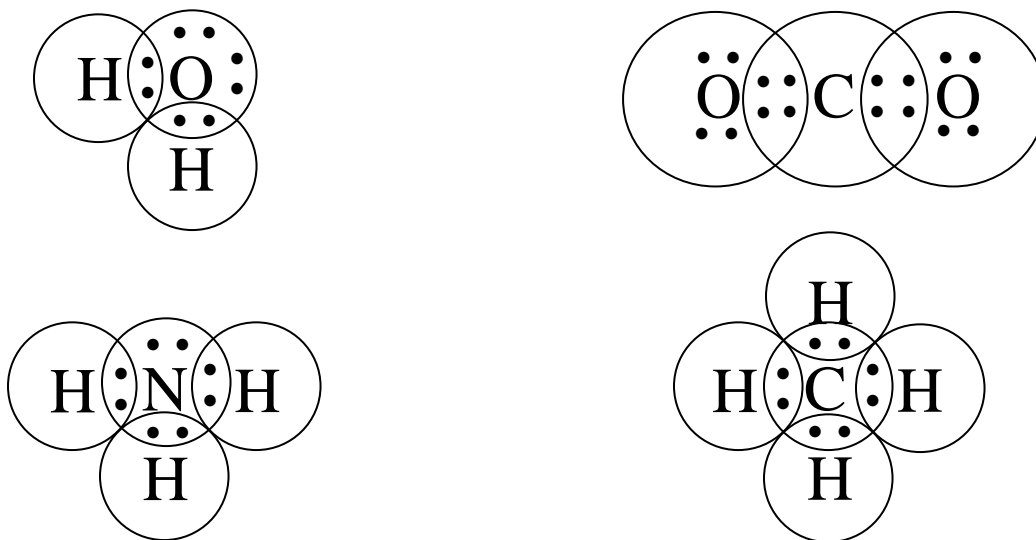


An electrostatic attraction increases as two hydrogen atoms approach each other. An energy minimum of -436 kJ/mole occurs at the internuclear distance of 0.074 nm . At this distance, the attractive forces balance the repulsive forces. The two “ $1s$ ” orbitals overlap and the two electrons interact equally with both nuclei. *The hydrogen molecule is at its most stable.*

A) Lewis Dot Models for Simple Molecules:

- In 1916, G. N. Lewis proposed that, by sharing electrons, two atoms could acquire a noble gas electron configuration.
- When a covalent bond is formed between two atoms, unpaired electrons from each atom usually pair up in such a manner so that the valence of each atom is filled.





B) Covalent Bond Lengths:

- When two atoms of the same element are held by a single covalent bond, the bond distance between the two atoms is approximately equal to the sum of their atomic radii.
- When two atoms of different elements are held by a single covalent bond, the bond distance between the two atoms is a bit less than the sum of their atomic radii. The more polar the bond, the shorter the bond distance is.

F–F 143 pm (atomic radius is 71 pm)

H–H 74 pm (atomic radius is 37 pm)

I–I 266 pm (atomic radius is 133pm)

H–F (more polar) 92 pm (sum of atomic radii is 108 pm)

H–I (less polar) 161 pm (sum of atomic radii is 170 pm)

- **Bond order** is the number of electron pairs being shared between any two bonded atoms.
- As the **bond order** between adjacent atoms increases, the length of the bonds decreases.

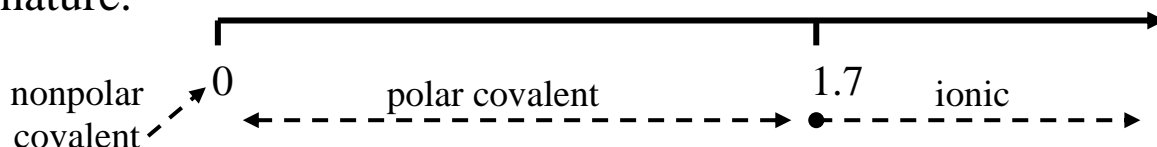
bond order = 1 C–C 154 pm (atomic radius is 77 pm)

bond order = 2 C=C 134 pm

bond order = 3 C≡C 120 pm

C) Bond Polarity and Electronegativity: [see figure 8.6 in text]

- **Electronegativity** is a measure of the ability of an atom in a molecule to attract electrons. The higher the electronegativity value of an atom, the more it will attract electrons to itself.
- A **nonpolar covalent bond** is one in which two atoms share electrons equally. As such, the electron density is symmetrical around the two atoms.
- A **polar covalent bond** is one in which two atoms share electrons unequally. As such, the electron density is asymmetrical around the two atoms. The electron density is greater around the atom with the greater electronegativity [*If the electron density becomes too unbalanced, the bond is characterized as **ionic***].
- The electronegativity difference can be used to predict bonding nature.



- Percent **ionic character** can also be used to predict bonding nature.

$$\%IC = \frac{EN_{hi} - EN_{lo}}{EN_{hi}} \times 100$$

A bond is characterized as nonpolar covalent if it has less than 5% ionic character. A bond is characterized as polar covalent if it has at least 5% but less than 50% ionic character. A bond is characterized as ionic if it has at least 50% ionic character.

Q2: Calculate the percent ionic character of NaF and HF and predict the type of bonding in each. [$EN_{Na} = 0.9$, $EN_F = 4.0$, and $EN_H = 2.1$]

<u>A2</u> : $\%IC \text{ in NaF} = \frac{EN_{hi} - EN_{lo}}{EN_{hi}} \times 100$	$\%IC \text{ in HF} = \frac{EN_{hi} - EN_{lo}}{EN_{hi}} \times 100$
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$$\%IC \text{ in NaF} = \frac{4.0 - 0.9}{4.0} \times 100$$

$$\%IC \text{ in HF} = \frac{4.0 - 2.1}{4.0} \times 100$$

$\%IC \text{ in NaF} = 78\%$
ionic bonding

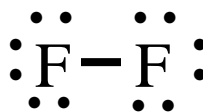
$\%IC \text{ in HF} = 48\%$
polar covalent bonding

D) Dipole Moments:

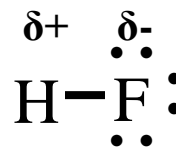
- A **nonpolar molecule** is one in which the electron density is symmetrically distributed over the entire molecule.
- A **polar molecule** is one in which the electron density is asymmetrically distributed over the entire molecule.



dipole moment = 0 D



dipole moment = 0 D



dipole moment = 1.82 D

- **Dipole moment** (μ) measures the degree of a molecule's polarity. The SI unit for dipole moment is the debye (D). 1.00 D is equal to 3.34×10^{-30} coulomb-meters (C-m).
- The larger the magnitude of the dipole moment is, the more polar the molecule.
- The magnitude of the **dipole moment** (μ) is the product of the magnitude of the charge (Q) and the distance of the separation of the charge (r).

$$\mu = Qr$$

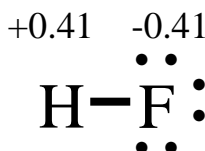
Q3: Calculate the effective charges (Q) on the H and F atoms of the HF molecule in unit of electron charge (e^-).

$$[\mu = 1.82 \text{ D}, \quad r = 92 \text{ pm}, \text{ and } 1 e^- = 1.60 \times 10^{-19} \text{ C}]$$

A3: $Q = \mu/r$

$$Q = \frac{1.82 \text{ D}}{92 \text{ pm}} \times \frac{1 \text{ pm}}{1 \times 10^{-12} \text{ m}} \times \frac{3.34 \times 10^{-30} \text{ C-m}}{1 \text{ D}} \times \frac{1 e^-}{1.60 \times 10^{-19} \text{ C}}$$

$$Q = 0.41 e^-$$



Compound	Bond Length	% IC	μ	Q
HF	92 pm	48%	1.82 D	0.41 e^-
HCl	127 pm	30.0%	1.08 D	0.178 e^-
HBr	141 pm	25%	0.82 D	0.12 e^-
HI	161 pm	16%	0.44 D	0.057 e^-

Q4: Using the Periodic Table, rank the bonds in each set in order of decreasing bond length and bond strength:

(i) S-F, S-Br, and S-Cl (ii) C=O, C-O, and C≡O

A4: (i) Bond length: S-Br > S-Cl > S-F

Bond strength: S-F > S-Cl > S-Br

(ii) Bond length: C-O > C=O > C≡O

Bond strength: C≡O > C=O > C-O

Q5: Arrange each set of bonds in order of increasing polarity, and indicate bond polarity with symbols:

(i) Cl-F, Br-Cl, and Cl-Cl

(ii) Si-Cl, P-Cl, S-Cl, and Si-Si

A5: (i) $\text{Cl}-\text{Cl} < \overset{\delta+}{\text{Br}}-\overset{\delta-}{\text{Cl}} < \overset{\delta+}{\text{Cl}}-\overset{\delta-}{\text{F}}$

(ii) $\text{Si}-\text{Si} < \overset{\delta+}{\text{S}}-\overset{\delta-}{\text{Cl}} < \overset{\delta+}{\text{P}}-\overset{\delta-}{\text{Cl}} < \overset{\delta+}{\text{Si}}-\overset{\delta-}{\text{Cl}}$

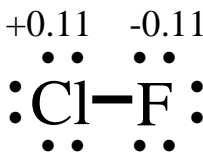
Q6: Calculate the value of Q (in e⁻) for a molecule of ClF.

[μ = 0.88 D, r = 163 pm, and 1 e⁻ = 1.60×10⁻¹⁹ C]

A6: Q = μ/r

$$Q = \frac{0.88 \text{ D}}{163 \text{ pm}} \times \frac{1 \text{ pm}}{1 \times 10^{-12} \text{ m}} \times \frac{3.34 \times 10^{-30} \text{ C-m}}{1 \text{ D}} \times \frac{1 \text{ e}^-}{1.60 \times 10^{-19} \text{ C}}$$

$$Q = 0.11 \text{ e}^-$$



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Lesson-3

IV – Lewis Models and Resonance Structures

A) Writing Lewis Models that Obey the Octet Rule:

- **Steps to determine a molecule's Lewis Model:**

- 1) Count the total number of valence electrons present in the **entire** formula.
- 2) From the formula...
 - (a) identify the central atom
 - (b) identify the peripheral atom(s).
- 3) Place the peripheral atom(s) as far from each other as possible around the central atom.
- 4) Connect the central atom to the peripheral atoms with a single (**sigma**) covalent bond.
- 5) Count the number of electrons used to bond the central atom to the peripheral atoms. If there are unused valence electrons, distribute them **evenly in pairs** around the peripheral atoms.
- 6) If electrons remain, place them on the central atom in pairs as nonbonding (“lone pair”) electrons.
- 7) If necessary, bring down one or more lone pair electrons from the peripheral atoms to bring the valence of the central atom to eight. This will create a multiple (**pi**) covalent bond.

Q1: H₂O

Q2: CCl_4

Q3: ClO_3^{1-}

Q4: HOCl

Q5: PCl_3

Q6: CO_2

Q7: HCHO

Q8: C_2H_2

Q9: O_2

B) Writing Lewis Models with Resonance Forms

- In some cases, a single Lewis model can not accurately depict a molecule or polyatomic ion. In these cases, more than one Lewis model is required and these are called **resonance forms**. Resonance is the result of electron-pair delocalization between several atoms.

Q10: CO_3^{2-}

Q11: SO_2

Q12: O_3

Q13: C_6H_6 (ringed structure)

Q14: HCO_2^{1-}

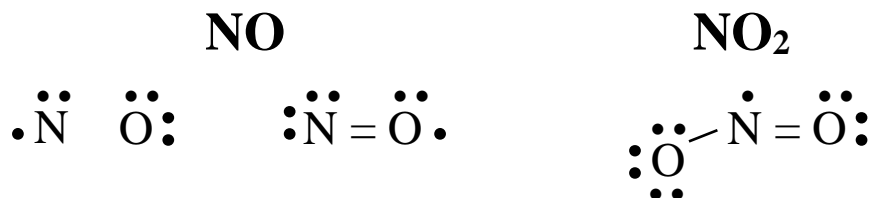
Q15: SO_3

Q16: P_2O_4

V – Exceptions to the Octet Rule

- **Formulas with Odd Numbers of Electrons:**

- Formulas with odd numbers of electrons (free radicals) can never satisfy the octet rule.



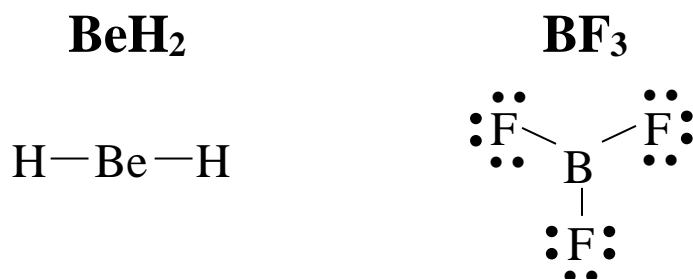
- **Oxygen Gas:**

- No Lewis structure of oxygen gas can satisfactorily explain the observed paramagnetism and double bond of this substance.

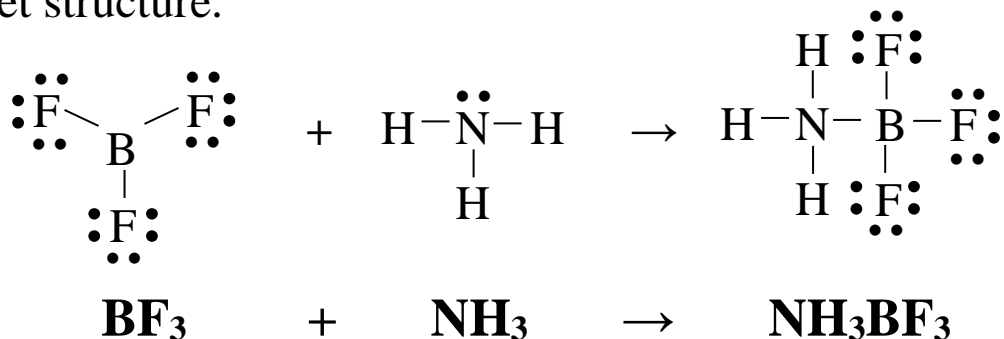


- **Formulas with Fewer than Eight Valence Electrons:**

- Molecules of beryllium and boron form compounds in which the central atoms have fewer than eight valence electrons.



- Reactions involving BF_3 are consistent with the incomplete octet structure.



In this reaction, a stable coordinate covalent bond is formed between the boron trifluoride molecule and the ammonia molecule.

- **Formulas with More than Eight Valence Electrons:**
 - Molecules and ions of heavier nonmetals (I, S, P, Xe, etc.) form structures in which the central atoms have more than eight valence electrons called an *expanded octet*.
 - Only elements in period 3 and beyond can act as the central atom in formulas that form structures with expanded octets.

Q17: ICl_4^{1-}

Q18: PCl_5

Q19: SF_6

Q20: IF_5

Q21: AsF_4^{1-}

Q22: XeF_2

Q23: ClF_3

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Lesson-4

VI – Oxidation Number and Formal Charge

A) Oxidation Number:

- In a formula, an atom's **oxidation number** is the ionic charge it would have if **ALL** the bonding electrons were completely transferred to the atom with the higher electronegativity.
- In order to determine each atom's oxidation number in this way, the formula's Lewis model must be drawn.
- The sum of the oxidation numbers for the atoms in a molecule or polyatomic ion must equal the total charge of the molecule or polyatomic ion.

Q1: Determine the oxidation numbers for each element in the following formulas: (a) HCl (b) CO₂ (c) HOCl

A1: (a) $\text{H}-\ddot{\text{Cl}}:$

Chlorine being more electronegative would have 8 valence electrons and hydrogen would have none.

Cl's ox# is -1 and H's ox# is +1.

(b) $\ddot{\text{O}}=\text{C}=\ddot{\text{O}}$

Oxygen being more electronegative would have 8 valence electrons and carbon would have none.

O's ox# is -2 and C's ox# is +4.

(c) $\text{H}-\ddot{\text{O}}-\ddot{\text{Cl}}:$

Oxygen being most electronegative would have 8 valence electrons, chlorine would have 6, and hydrogen would have none.

O's ox# is -2, Cl's ox is +1, and H's ox# is +1.

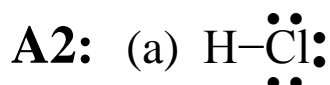
B) Formal Charge:

- In a formula, an atom's formal charge is the charge it would have if the **ALL** the bonding electrons were shared equally.
- In order to determine each atom's formal charge in this way, the formula's Lewis model must be drawn.
- The following equation can be used to determine an atom's formal charge:

$$\text{formal charge} = \# \text{ valence } e^- - [\# \text{ unshared } e^- + \frac{1}{2}\# \text{ shared } e^-]$$

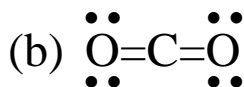
- The sum of the formal charges for the atoms in a molecule or polyatomic ion must equal the total charge of the molecule or polyatomic ion.

Q2: Determine the formal charges for each element in the following formulas: (a) HCl (b) CO₂ (c) HOCl



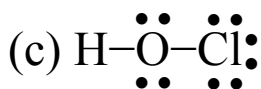
Hydrogen would have 1 valence electron and chlorine would have 7.

Cl's formal charge is 0 and H's formal charge is 0.



Carbon would have 4 valence electrons and each of the oxygen atoms would have 6.

C's formal charge is 0 and O's formal charge is 0.



Hydrogen would have 1 valence electron, oxygen would have 6, and chlorine would have 7.

H's formal charge is 0, O's formal charge is 0, and Cl's formal charge is 0.

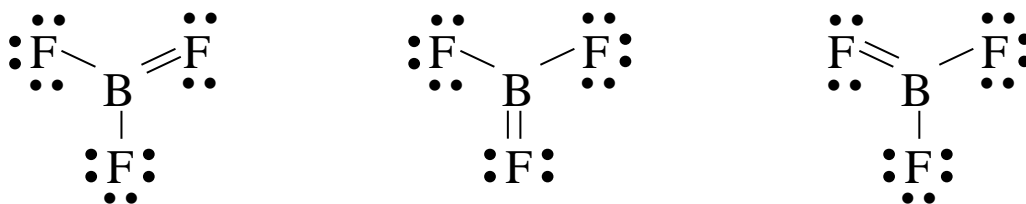
Remember, that oxidation numbers and formal charges do NOT actually exist. They are each a "bookkeeping" system that are used to monitor electrons.

C) Selecting the Most Prominent Resonance Structure:

- Formal charge can be used to select the most prominent or best resonance form of a formula.
 - Smaller formal charges are preferable to larger ones.**
 - The atom with a higher EN value should have the more negative formal charge.**

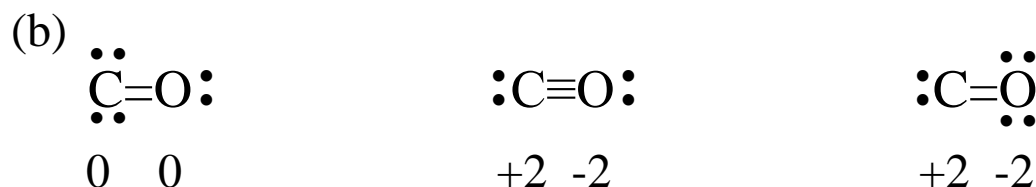
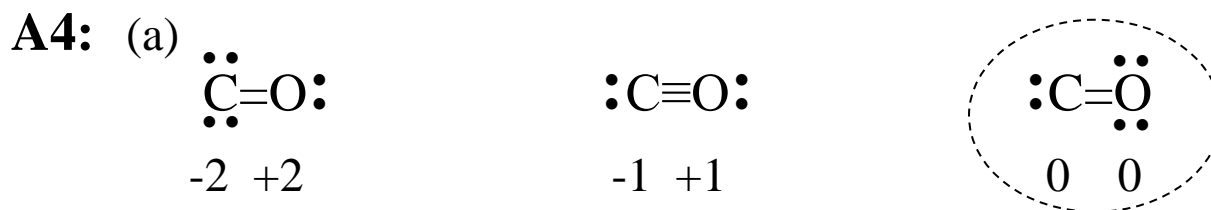
Q3: In terms of Lewis structures and formal charges, explain why the boron atom in BF_3 does not complete the octet

A3:



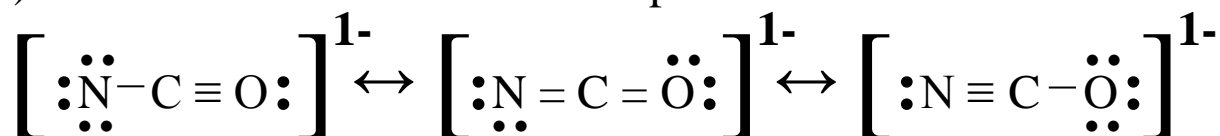
If the octet were completed around boron in BF_3 , then three equal resonance structures would be required and the fluorine in the $\text{B}=\text{F}$ bond would have a +1 formal charge. YIKES!

Q4: Carbon monoxide has three resonance Lewis structures, two of which violate the octet rule in that one atom has fewer than eight electrons. (a) Draw the three Lewis structures and use formal charge to determine the most prominent structure. (b) Determine the oxidation number for each atom in each structure.

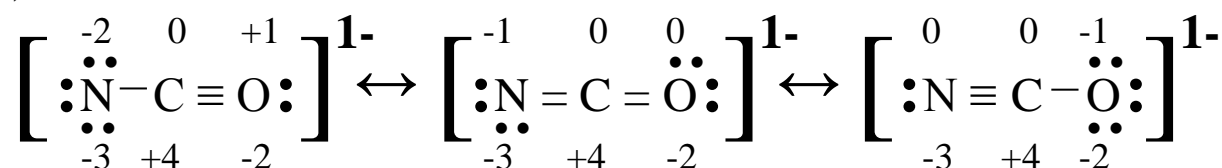


Q5: The following are three possible Lewis models for the OCN^{1-} ion.

- (a) Determine the formal charges [above each atom] and the oxidation numbers [below each atom] for each atom in each of the structures.
 (b) Which Lewis model is the most preferred one? WHY?
 (c) Which Lewis model is the least preferred one? WHY?



A5: (a)

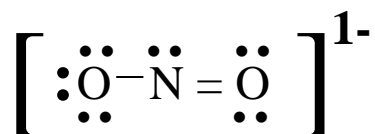


- (b) The Lewis model on the right is the most preferred one, because the oxygen, the most electronegative element, is assigned the formal charge of -1.
 (c) The Lewis model on the left is the least preferred one, because the oxygen, the most electronegative element, is assigned the formal charge of +1.

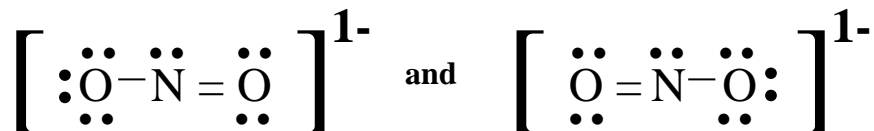
Q6: (a) Write the Lewis model for the nitrite ion (NO_2^{1-}).

- (b) Are resonance structures needed to describe the nitrite ion?
 If so, write all Lewis models.
 (c) Based on your Lewis model, predict the bond order of NO_2^{1-} ?

A6: (a)



(b) Yes.



(c) bond order = 1 

Chapter-8

Lesson-5

VII – Bond Enthalpy

- **Bond enthalpy** is the energy **ABSORBED** in order to break one mole of bonds in the gaseous phase (kJ/mol). Bond enthalpy is always a positive quantity. The designation D(bond type) is used to represent bond enthalpy.



This process is called atomization.

- When one mole of bonds is formed energy is released and is equal to the negative of the **bond enthalpy**.

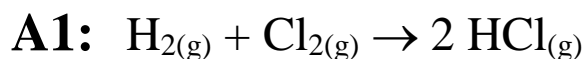


- **Bond enthalpy** can be used to estimate the ΔH_R for a reaction.

$$\Delta H_R = \sum D(\text{bonds broken}) - \sum D(\text{bonds formed})$$

Caution: This is not the same as $\Delta H_R = \sum H_f \text{ of products} - \sum H_f \text{ of reactants}$

Q1: Using bond enthalpies, estimate the ΔH_R for the formation of HCl from its free elements.



$$\Delta H_R = [D(\text{H-H}) + D(\text{Cl-Cl})] - [2D(\text{H-Cl})]$$

$$\Delta H_R = [D(436) + D(242)] - [2D(431)]$$

$$\boxed{\Delta H_R = -184 \text{ kJ}} \text{ this compares well to the } \Delta H_f \text{ of HCl on Appendix C.}$$

Q2: Using bond enthalpies, estimate the ΔH_R for the formation of ICl from its free elements. $D(I_2) = 151 \text{ kJ/mol}$, $D(Cl_2) = 242 \text{ kJ/mol}$, $D(I-Cl) = 208 \text{ kJ/mol}$

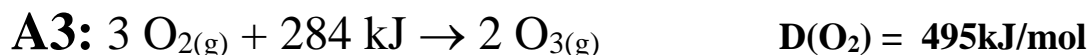


$$\Delta H_R = [D(I-I) + D(Cl-Cl)] - [2D(I-Cl)]$$

$$\Delta H_R = [D(151) + D(242)] - [2D(208)]$$

$$\Delta H_R = -23 \text{ kJ}$$

Q3: Estimate the bond enthalpy of the $1\frac{1}{2}$ resonance bond that exists between each oxygen atom in a molecule of ozone (O_3).



$$\Delta H_R = [3D(O=O)] - [4D(O\cdot O)]$$

$$284 = [3D(495)] - [4D(O\cdot O)]$$

$$D(O\cdot O) = 300 \text{ kJ/mol} \quad \text{this makes sense since the } D(O-O) = 146 \text{ kJ/mol}$$

A) Strengths of Covalent Bonds:

- A covalent bond with a large bond enthalpy tends to be stronger (and more stable) than one with a smaller bond enthalpy.
- **Bond order** is the number of electron pairs being shared between any two bonded atoms.
- The **bond enthalpy** is greater for the same two atoms when the bond order is greater.



- If $N_{2(g)} + 3 H_{2(g)} \rightarrow 2 NH_{3(g)} + 92 \text{ kJ}$ then what is $D(N_2)$?

$$\Delta H_f = [D(N_2) + 3 D(H-H)] - [6D(N-H)]$$

$$-92 = [D(N_2) + 3(436)] - [6(391)]$$

$$D(N_2) = 946 \text{ kJ/mol} \quad \text{this confirms the bond order of } N_2!$$

B) Bond Enthalpy and Bond Length:

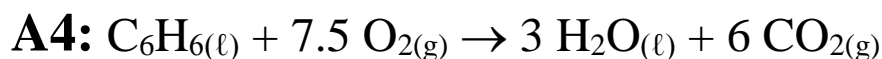
- As the bond order between two atoms increases, the length of the bond decreases.
- As the length of the bond between two adjacent atoms decreases, the strength of the bond between the atoms generally increases.

C–C	154 pm	348 kJ/mol
C=C	134 pm	614 kJ/mol
C≡C	120 pm	839 kJ/mol

- In general, the more polar a bond is the stronger it will be in comparison to the average bond enthalpy of the two bonding atoms.

H–H (37 pm for H)	74 pm	436 kJ/mol
F–F (72 pm for F)	143 pm	153 kJ/mol
I–I (133 pm for I)	266 pm	151 kJ/mol
H–F (more polar)	92 pm*	565 kJ/mol*
H–I (less polar)	161 pm*	297 kJ/mol*

Q4: If ΔH_R for the combustion of benzene (C_6H_6) is -3267 kJ/mol, $D(C-H)$ is +413 kJ/mol, $D(O_2)$ is +495, $D(H-O)$ is +463 kJ/mol, and $D(C=O)$ is 799 kJ/mol then calculate $D(C_2)$ and determine the carbon to carbon bond order.

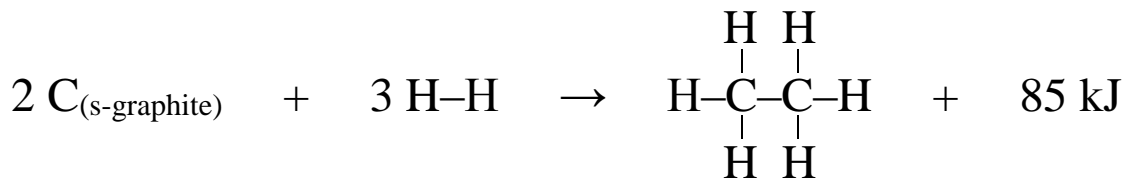


$$\Delta H_f = [6D(C_2) + 6D(C-H) + 7.5 D(O_2)] - [6D(H-O) + 12D(C=O)]$$
$$-3267 = [6D(C_2) + 6(413) + 7.5(495)] - [6(463) + 12(799)]$$

$$D(C_2) = +485 \text{ kJ/mol}$$

This value is very close to the average of $D(C-C)$ and $D(C=C)$. This suggests a bond order of $1\frac{1}{2}$ which is consistent with the carbon to carbon resonance bonding found in the benzene molecule.

Q5: If ΔH_R for the formation of ethane (C_2H_6) is -85 kJ/mol , and if $D(H-H)$ is $+436 \text{ kJ/mol}$, $D(C-C)$ is $+348 \text{ kJ/mol}$, and $D(C-H)$ is $+413 \text{ kJ/mol}$, then calculate $D(C_{(s\text{-graphite})})$ base on the reaction equation below.

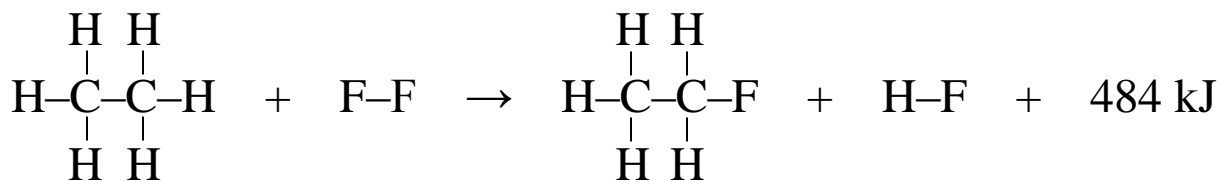


A5: $\Delta H_f = [2D(C_{(s\text{-graphite})}) + 3D(H-H)] - [1D(C-C) + 6D(C-H)]$

$$-85 = [2D(X) + 3D(436)] - [1D(348) + 6D(413)]$$

$$X = 717 \text{ kJ/mol}$$

Q6: Ethane gas reacts with fluorine gas according to the following equation:



If $D(C-C)$ is $+348 \text{ kJ/mol}$, $D(C-H)$ is $+413 \text{ kJ/mol}$, $D(F-F)$ is $+153 \text{ kJ/mol}$, and $D(H-F)$ is $+565 \text{ kJ/mol}$, then calculate $D(C-F)$.

A6: $\Delta H_f = [1D(C-C) + 6D(C-H) + 1D(F-F)] -$
 $[1D(C-C) + 5D(C-H) + 1D(C-F) + 1D(H-F)]$

$$-484 = [1D(348) + 6D(413) + 1D(153)] -$$

$$[1D(348) + 5D(413) + 1D(X) + 1D(565)]$$

$$X = 485 \text{ kJ/mol}$$